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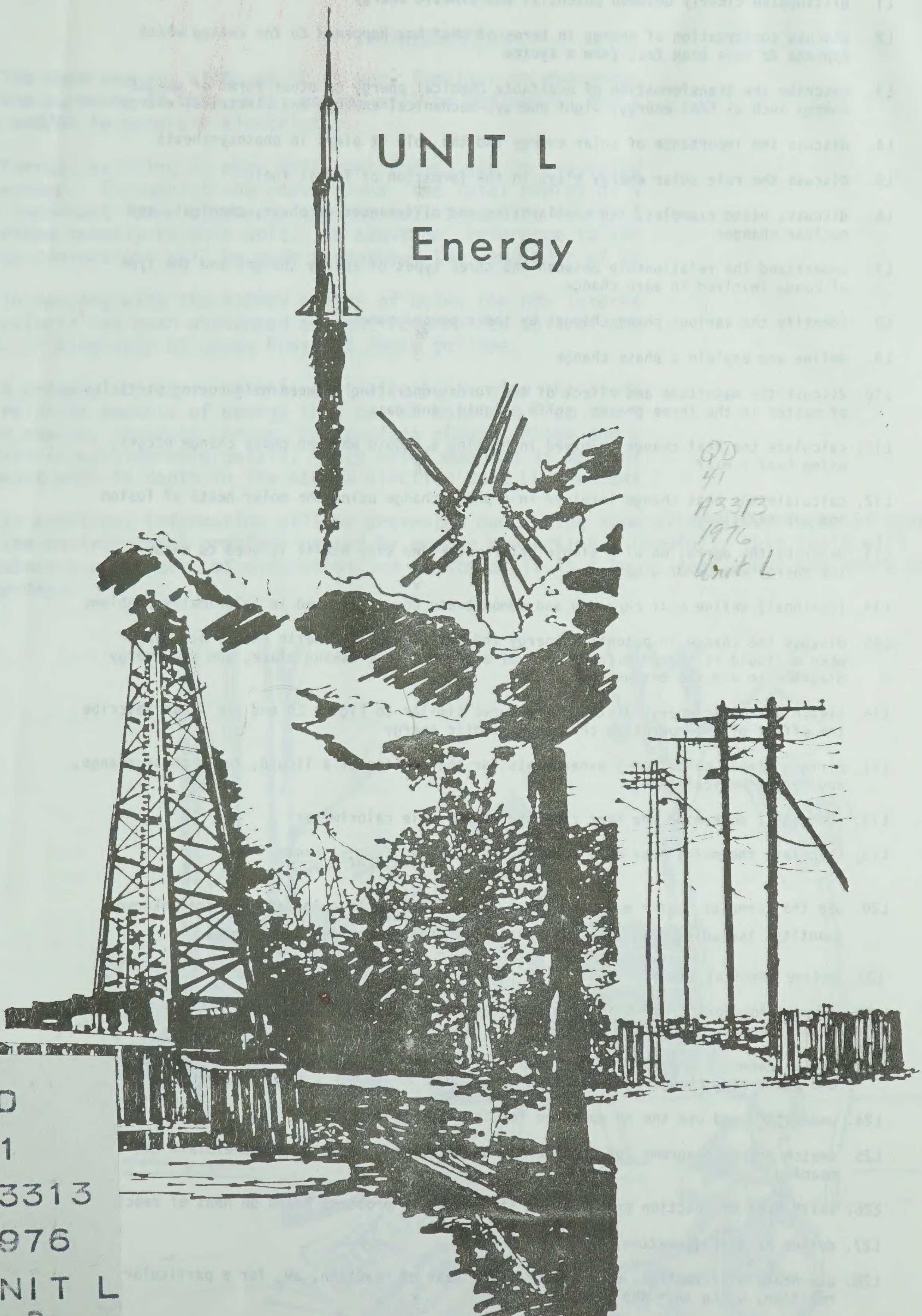
UNIT L

Energy



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Unit L

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OBJECTIVES

L, *Energy*, the student should be able to:

L1. distinguish clearly between potential and kinetic energy

L2. discuss conservation of energy in terms of *what has happened to the energy which appears to have been lost from a system*

L3. describe the transformation of *available* chemical energy to other forms of *useful* energy such as heat energy, light energy, mechanical energy, and electrical energy

L4. discuss the importance of solar energy and the role it plays in photosynthesis

L5. discuss the role solar energy plays in the formation of fossil fuels

L6. discuss, using examples, the similarities and differences of phase, chemical, and nuclear changes

L7. understand the relationship between the three types of energy changes and the type of *bonds* involved in each change

L8. identify the various phase changes by their proper names

L9. define and explain a phase change

L10. discuss the magnitude and effect of the forces operating between neighboring particles of matter in the three phases, solid, liquid, and gas

L11. calculate the heat change involved in heating a liquid when no phase change occurs, using $heat = ms\Delta t$

L12. calculate the heat change involved in a phase change using the molar heats of fusion and vaporization

L13. describe the operation of a simple calorimeter and show how it is used to measure the energy change of a system

L14. (Optional) define *heat capacity* and demonstrate how it is used in calorimetry problems

L15. discuss the change in potential energy and the change in kinetic energy occuring when a liquid is heated *with* or *without* a phase change taking place, and use energy diagrams to aid the discussion

L16. sketch a kinetic energy distribution curve similar to Figure L5 and use it to describe the effect of a temperature change on kinetic energy

L17. perform simple calorimetry experiments for the heating of a liquid, for a phase change, and for a chemical change

L18. (Optional) determine the heat capacity of a simple calorimeter

L19. calculate the *molar heat* of a substance using $molar\ heat = \frac{heat}{\#moles}$

L20. use the formulas $heat = ms\Delta t$ and $molar\ heat = \frac{heat}{\#\ moles}$ to calculate any one unknown quantity, including mass, specific heat, temperature change, and # moles

L21. define chemical change

L22. discuss four practical applications of *exothermic* chemical reactions using a different form of energy each time

L23. define *chemical bond* and *bond energy* and show the relationship between *bond energy* and *net energy change of a reaction*

L24. understand and use the ΔH *notation* for the *heat of reaction*

L25 sketch energy diagrams for endothermic and exothermic reactions and explain their meaning

L26. solve heat of reaction problems and stoichiometry problems based on heat of reaction

L27. define *heat of formation*, H

L28. use heats of formation, H, to predict the heat of reaction, ΔH, for a particular reaction, using $\Delta H = \Sigma H_p - \Sigma H_R$

L29. describe briefly a nuclear change

L30. compare phase, chemical, and nuclear changes in terms of the products produced, and, in terms of the *relative* amount of energy associated with each type of change

L31. discuss the topic *alternate forms of energy* and the environmental problems associated with the various methods of producing energy

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ENERGY

INTRODUCTION

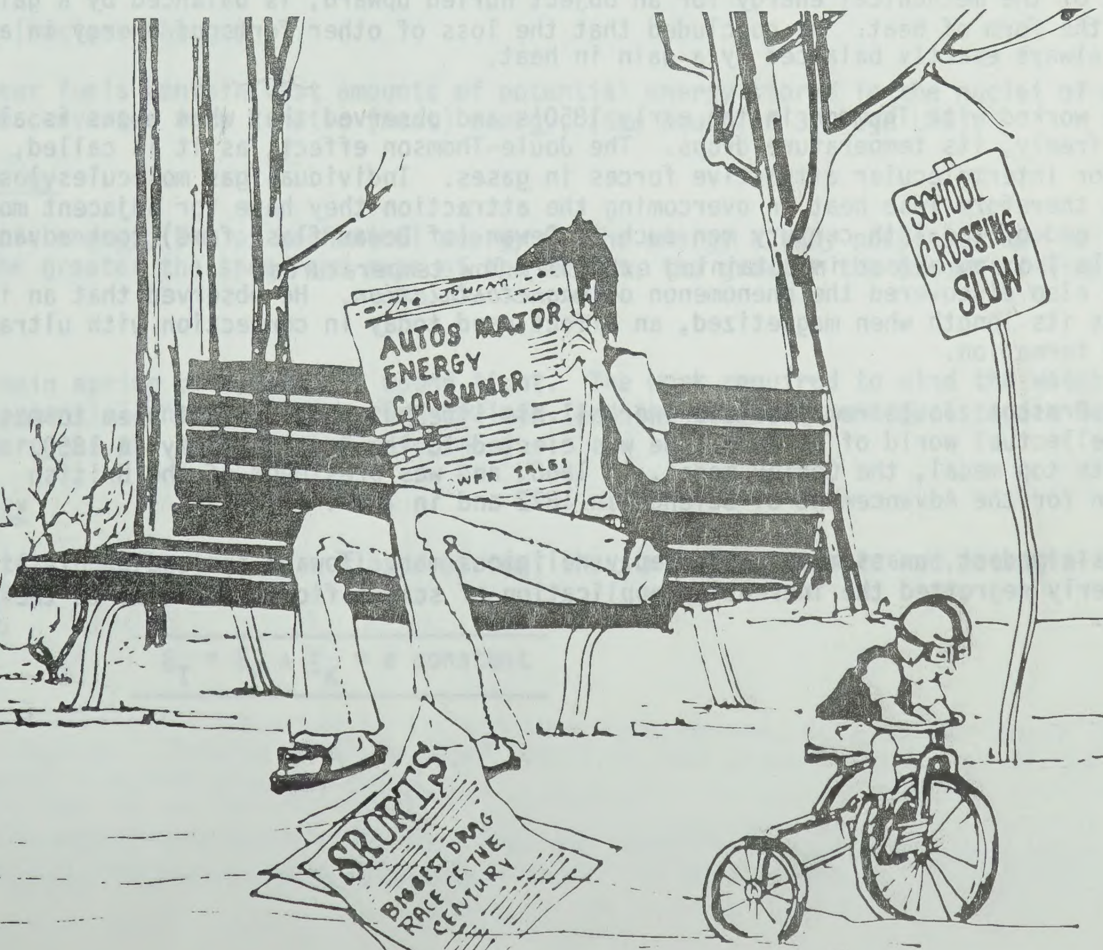
The term *energy*, although it is very familiar to everyone, is difficult to define. For the purposes of this unit energy will be defined as the capacity to do work, to create heat and/or to generate electricity.

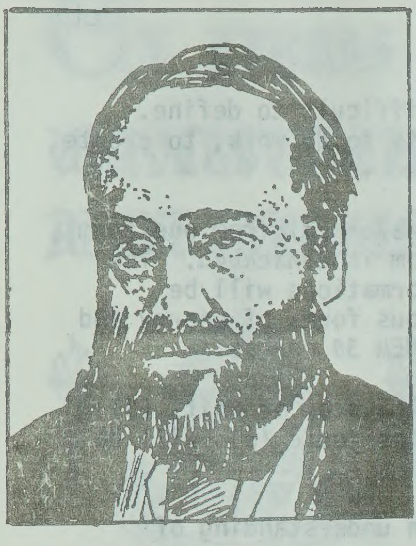
Energy, existing in many different forms, can be *converted* (*transformed*) from one form to another. Throughout the conversions the total energy of a system is *conserved*. The most important forms of energy and the most important energy transformations will be described briefly in this unit. In addition reference to the various forms of energy and energy conversions will be made throughout the remainder of the *ALCHEM 30* course.

In keeping with the *ALCHEM* policy of using the new International System of Units (SI), the calorie has been abandoned and the recommended unit of energy, the *joule* (J), will be used. A biography of James Prescott Joule follows.

A primary objective of this unit will be to give the students an understanding of the relative amounts of energy that can be obtained from the three types of energy changes- *phase change*, *chemical change*, and *nuclear change*. Phase and chemical changes will be studied in considerable detail. Only brief mention is made of nuclear change as this topic is dealt with in depth in the *ALCHEM* elective Unit LL, *Nuclear Energy and Radiation*.

In addition, information will be presented concerning some alternative forms of energy and the environmental problems caused by energy production. Hopefully this topic will stimulate a great deal of discussion and debate as it is a topic read about and heard about every day.





JAMES PRESCOTT JOULE

1818 - 1889

James Prescott Joule was the son of a brewery owner in Manchester, England. Coming from a very wealthy family, Joule was able to devote much of his time to study and research. He received some instruction from John Dalton but he was mostly self-educated.

Joule was a superb experimenter, almost fanatical about measuring the heat produced in just about any kind of process. In every case he calculated the work entering a system and the heat coming out, and in every case he found the two were closely related. After about 10 years of intensive experimentation and measurement Joule developed the most reliable value for *the mechanical equivalent of heat.* With the help of William Thomson (later known as Lord Kelvin) he was able to convince the science world to accept the results of his years of investigation and finally he received the credit he so justly deserved. A new unit of work, the joule (equivalent to a newton-metre), was named in his honor.

Joule found the amount of work required to raise of temperature of 1.00 g of water 1.00°C. This amount of work was equated to 4.19 joules (4.19 J) and was called the *mechanical equivalent of heat.*

Joule's determination of the mechanical equivalent of heat did much to enhance the notion of conservation of energy. He reasoned that the loss, through friction and air resistance of the mechanical energy for an object hurled upward, is balanced by a gain of energy in the form of heat. He concluded that the loss of other forms of energy in any system is always exactly balanced by a gain in heat.

Joule worked with Thomson in the early 1850's and observed that when a gas is allowed to expand freely, its temperature drops. The Joule-Thomson effect, as it is called, is evidence for intermolecular attractive forces in gases. Individual gas molecules lose energy and therefore lose heat in overcoming the attraction they have for adjacent molecules. Toward the end of the 19th century men such as Dewar (of Dewar flask fame) took advantage of the Joule-Thomson effect in obtaining extremely low temperatures.

Joule also discovered the phenomenon of *magnetostriction.* He observed that an iron bar changes its length when magnetized, an effect used today in connection with ultrasonic sound wave formation.

James Prescott Joule remained a brewer all his life but this did not seem to matter to the intellectual world of science. He was elected to the Royal Society in 1850, received its top medal, the Copley medal, in 1866, and was president of the British Association for the Advancement of Science in 1872 and in 1887.

He was a modest, unassuming, and deeply religious man. Toward the end of his life, Joule bitterly regretted the increasing application of scientific discoveries to the art of warfare.

FORMS OF ENERGY

Potential Energy

Potential energy (Ep) is the energy stored in a body due to its position. Potential energy is an inactive form of energy but the potential to produce usable energy is there if properly activated.

Note:

As the next section deals with energy changes, the four examples below describe the conversion of available potential energy to other usable forms of energy by means of phase, chemical, and nuclear changes.

Example 1: (Phase Change)

The potential energy stored in *steam* may be used to heat buildings. (See Example 2, Page L4.)

Example 2: (Chemical Change)

Fossil fuels such as coal, oil, and natural gas all possess potential energy in a *chemical form* because heat (kinetic energy) and light (potential energy) are produced if they are burned in a chemical reaction.

Example 3: (Chemical Change)

A battery stores potential energy in *chemical form* which can be easily and conveniently converted into electrical energy upon demand.
(See *Electrochemistry*, ALCHEM 30, Unit M.)

Example 4: (Nuclear Change)

Nuclear fuels contain vast amounts of potential energy stored in the nuclei of atoms that can be converted into kinetic (heat) energy. (See Equation 3, Page L36).

Kinetic Energy

Kinetic energy (Ek) or mechanical energy is the energy a body possesses due to its motion. The greater the speed and mass of the object, the greater the amount of kinetic energy.

Example:

The main spring of a watch is wound tight. The work required to wind the watch is now stored as potential energy. The spring unwinds, and the potential energy is again converted to kinetic energy as the hands of the watch move.

Total Energy

The total energy (E_T) in a closed system, the sum of the kinetic and potential energies, is a constant.

$$E_T = E_P + E_K = \text{a constant}$$

CHEMICAL AND NUCLEAR CHANGES

A *phase change* is a change in state of matter such as the melting of a solid and the vaporization of a liquid. During a phase change no new chemical species is produced, but the new phase that is produced has a different amount of energy than the old phase. Matter either releases energy or absorbs energy during a phase change.

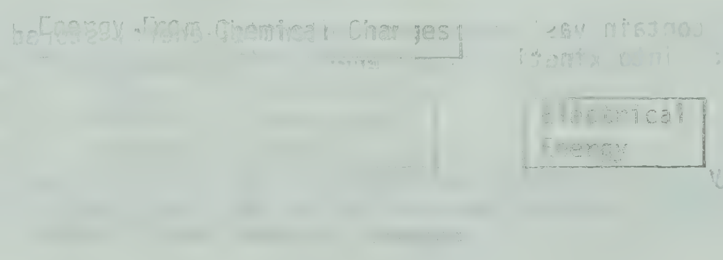
Example:

Large building complexes consisting of several buildings sometimes are heated by one large centrally located steam generating plant. The steam is piped into the heating systems of the various buildings where it is used either to heat the air or to heat water which is then piped throughout the building to provide *hot water heating*. The University of Alberta, Edmonton campus, uses this form of *hot water heating*. The amount of heat released when steam condenses will be discussed later in the section *Energy Associated With a Phase Change*.

Chemical Change

A *chemical change* is a change in which one substance, with a certain amount of energy, is converted into another substance with a different amount of energy.

Usable forms of energy can be obtained either directly or indirectly from chemical changes (see Figure L1



Example 1: (Conversion of Chemical Energy into Heat Energy)

Fossil fuels such as coal, oil, and natural gas all possess *potential energy* in a *chemical form* because heat is produced if they are burned in a chemical reaction. For example, most homes in Alberta are heated with natural gas.

For centuries heat energy was used for heating and cooking. During the 18th and 19th centuries heat energy was used to generate steam to operate turbines or engines. For the past 50 years an ever increasing amount of total heat energy has been used to generate electrical energy.

Example 2: (Conversion of Chemical Energy into Electrical Energy) Unit M

Electrical energy is a special intermediate form of energy that can be easily and economically converted into kinetic energy or directly into heat or light energy. ALCHEM, *Electrochemistry*, discusses the conversion of chemical energy into electrical energy and its subsequent practical applications ([illegible]).

PHASE, CHEMICAL AND NUCLEAR CHANGES

Example 3: (Conversion of Chemical Energy into Light Energy)

Chemical energy usually releases energy in the form of heat or mechanical energy, but in a few reactions it is released directly as light energy. Phosphorus or diphosphorus trioxide produces light when it is oxidized by air. A second example is the cold light of certain living organisms such as the fire-fly. The light of the fire-fly is produced by the oxidation of a protein derivative, luciferin, within the fly. In this reaction, light energy is *produced* by the chemical change. In other reactions such as *photosynthesis*, light energy is *absorbed* during the chemical change. Photosynthesis is discussed in the next section dealing with *Nuclear Change*.

Example 4: (Conversion of Chemical Energy into Mechanical Energy)

The conversion of chemical energy into mechanical energy is best exemplified by *explosions*, reactions in which large volumes of gases and large amounts of heat are released very quickly. The heat released causes the gases produced to expand very quickly, and this expansion can be used to do work. For example, the explosion of gunpowder in a rifle produces rapidly expanding gases which propel the bullet out of the barrel. Similarly, the explosion of gasoline in an internal combustion engine produces rapidly expanding gases which drive the pistons of the automobile engine.

Note:

The automobile engine is an extremely inefficient engine as 85% of the chemical energy of the gasoline is wasted, leaving only 15% to drive the car. Approximately 70% of the energy is wasted in the form of heat either as exhaust gases or by absorption by the water and oil. About 5% is lost due to engine friction, and a further 10% is wasted through transmission losses.

Nuclear Change

A *nuclear change* is a change in which one element is converted into another element (or elements), and this change is accompanied by the absorption or release of extremely large quantities of energy.

Example 1:

Fantastic amounts of energy are stored as potential energy in the nuclei of atoms. In a nuclear reactor the nuclei are disrupted and the tremendous binding energy of the nucleons is suddenly released in an enormous exothermic reaction. The staggering amount of energy available from nuclear reactors will be discussed in the *ALCHEM* elective Unit LL.

Large quantities of energy are produced in nuclear reactors by *exothermic* nuclear reactions. However, scientists use *endothermic* nuclear reactions to produce the heavy elements and to synthesize new elements.

Example 2:

Nuclear reactions, occurring in the atmosphere of the sun, produce the energy of the sun, *solar energy*. (Details of the nuclear reactions that produce the energy of the sun are discussed in *ALCHEM* elective Unit LL.)

Solar energy (light energy from solar nuclear reactions), causes the most important of all chemical reactions, *photosynthesis*, to occur. In photosynthesis, green plants absorb

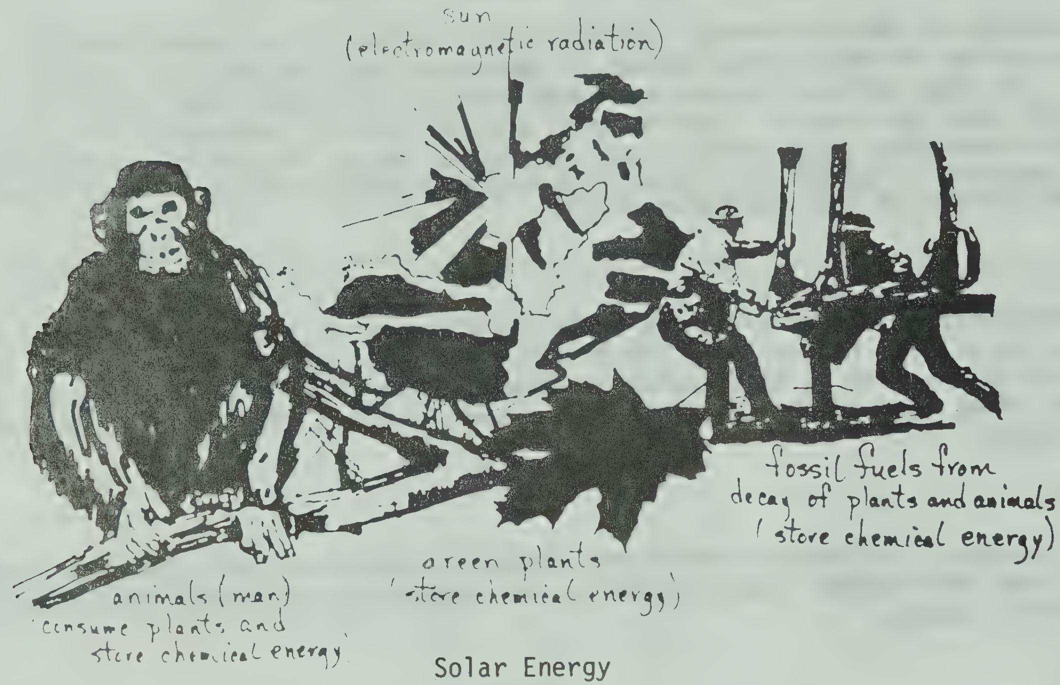
PHASE, CHEMICAL AND NUCLEAR CHANGES

sunlight and convert carbon dioxide and water vapor from the air into carbohydrates and oxygen. The reaction is catalyzed by chlorophyll, the green coloring matter of plants.

$$CO_{2(g)} + H_2O_{(g)} + \text{light} \xrightarrow{\text{chlorophyll}} \text{carbohydrates} + \text{oxygen}$$

$$CO_2(g) + H_2O(g) + 2824\ \text{kJ} \longrightarrow \underset{\text{glucose}}{C_6H_{12}O_6(s)} + 6O_2(g)$$

Glucose is not the end product of photosynthesis although it is found in plants. The plants form more complex carbohydrates from glucose, such as sucrose, starches and cellulose.



Solar Energy
Figure L2

Radiant energy from the sun is stored as chemical energy by plants and then by other living organisms. Gradually the chemical energy is stored in the fossil fuels used by man as the source of most of his energy (read *Solar Energy* page L41 before going on).

ENERGY

ENERGY CHANGES

Energy and Bonds

All particles of matter possess potential energy as a result of forces acting on them. The force of attraction among the particles in matter is called a *bond*. By breaking existing bonds and forming new bonds, energy can either be added to the matter or removed from the matter. The amount of energy involved and the kind of changes that occur depend on the type of bonds involved.

Example 1:

Consider the phase change, $H_2O_{(\ell)} + 41\ kJ \longrightarrow H_2O_{(g)}$. Energy is required by the water to change it to steam. The energy absorbed does work to overcome the *inter*molecular attractive forces bonding the water molecules together. The equation shows that the steam has 44 kJ more energy than the water.

Example 2:

Consider the chemical change $2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(\ell)} + 484\ kJ$

Energy is required to *break* two moles of H-H bonds and one mole of O-O bonds. Energy is released when four moles of H-O bonds are *formed*. The equation shows that 484 kJ more energy is released by *forming* bonds than was required to *break* bonds. (This concept will be expanded later.)

Table L1 shows how the type of changes and type of bonds are related.

Table L1
Energy Changes

	Type of Change	Type of Bond	Particles Involved	Fuel
1.	phase change	*inter*molecular bond (electostatic force)	molecules or ions	steam heat
2.	chemical change	*intra*molecular bond (electrostatic force)	atoms or ions	natural gas, gasoline, oil, coal, wood
3.	nuclear change	nuclear bond (nuclear force)	nucleons	radioactive decay material, Co fission material, $^{235}_{92}U$ fusion material, ^2_1H

This unit will look at the type of forces involved in each kind of change, write the equations for these changes and calculate the change in energy involved for each type of energy. (Nuclear change is covered more thoroughly in the elective *ALCHEM* Unit LL which is designed to accompany *ALCHEM* Unit L.)

The energy involved in forming and breaking bonds in any kind of ~~a~~ change can be summarized as follows:

bonded particles + energy ⟶ *separated particles* (old bonds broken)
separated particles ⟶ *bonded particles + energy* (new bonds formed)

In general, the amount of energy increases in the order; phase change, chemical change, nuclear change. The energy change is proportional to the strength of the bonds involved.

Nuclear changes produce or require energy in the amount of 10^6 to 10^9 kJ

PHASE, CHEMICAL AND NUCLEAR CHANGES

Questions

1. The recommended unit of energy in SI units is the ________________________.

2. What is meant by the term *mechanical equivalent of heat?*

3. Potential energy is defined as ____

____

4. Kinetic energy is defined as ____

____

5. Compare, using examples, *available* energy with *usable* energy.

6. Describe, with the aid of a chemical equation, the process of photosynthesis.

7. Discuss how the energy of fossil fuels such as coal can be traced back to the sun.

8. Give a practical example of each of the following.

   Phase Change—

   Chemical Change—

   Nuclear Change—

MEASUREMENT OF ENERGY CHANGE

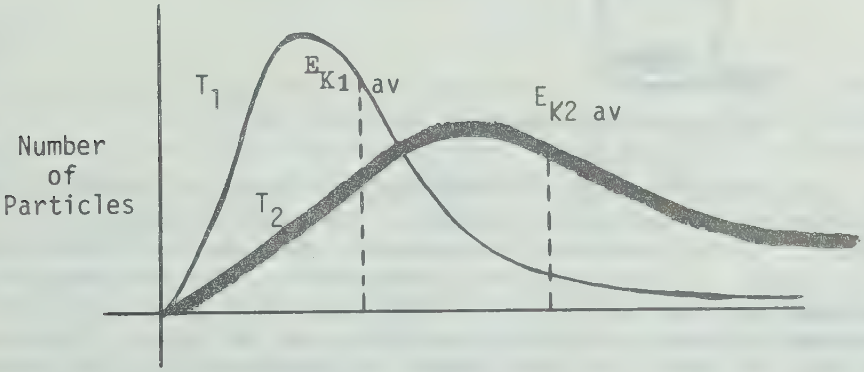
Affect of Temperature on Kinetic Energy

In a sample of any substance the particles move at various speeds and possess various kinetic energies. The range of energies shown in Figure L5 is a result of molecular collisions which constantly change the speed of a particular particle. Consider a sample of hydrogen gas. A particular fast moving hydrogen molecule collides with a slow moving molecule. The faster molecule transfers some of its energy to the slower molecule. The faster one slows down and the slower one speeds up. This exchange of kinetic energy is constantly going on but the total kinetic energy of the gas stays the same provided the temperature stays the same.

Temperature is a measure of the degree of hotness of a sample, and is proportional to the average kinetic energy of all the particles of the sample.

Note:
Heat and temperature do not refer to the same quantities.

If the temperature of a substance is raised, what happens to the particles? The heat energy added is absorbed by the particles thus increasing the average kinetic energy. Figure L5 shows the distribution of kinetic energy in a gas sample at two different temperatures, T_1 and T_2. The *heavier* curve represents the *higher* temperature, T_2.



Kinetic Energy, E_k
Distribution of Kinetic Energy at Two Temperatures
Figure L5

At the higher temperature, T_2, the particles have a higher average kinetic energy, $E_{K2\ av}$, than at the lower temperature. Thus temperature can be considered a measure of the average kinetic energy of the particles. Experimentation has verified that kinetic energy is directly proportional to absolute temperature. (Absolute temperature or Kelvin temperature is Celsius temperature plus 273.)

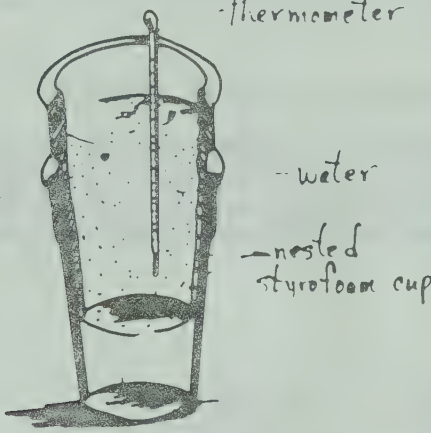
Figure L5 shows, that at the higher temperature, not only is the average kinetic energy higher, but there are *fewer* particles with *low* kinetic energy and *more* particles with *high* kinetic energy.

MEASUREMENT OF ENERGY CHANGE

Calorimetry

The *calorimeter* is a device used to *measure* the amount of energy emitted or absorbed during a change. The first simple calorimeter was devised and used by Antoine Lavoisier and Pierre Laplace during the early 1780's. They used the *mass of ice* melted as a measure of the amount of heat produced in the change. Today the temperature change produced in a given *mass of water* is used as a measure of the amount of heat produced or absorbed during a phase or chemical change.

A pyrex beaker, but preferably a styrofoam cup is generally used as a simple calorimeter in the high school laboratory.



A Simple Calorimeter
Figure L4

Figure L4 shows a typical calorimeter set-up. A known mass of water is placed in the calorimeter and the temperature of the water is recorded. The reacting substance is placed in the reaction vessel. If the reaction is exothermic the water will absorb heat and its temperature will *increase*. If the reaction is endothermic, the reacting substance will absorb heat from the water and the temperature of the water will *decrease*. The assumption is made that no heat is lost or gained by anything but the water. Knowing the *specific heat* of water the mass of water, and the change in temperature, it is a simple matter to calculate the energy change of the reaction, using the formula *heat* = $ms\Delta t$ where:

m is the mass in grams (g)
s is the specific heat in joules per gram degree Celsius (J/(g·°C))
Δt is the change in temperature in degrees Celsius (°C).

The specific heat is the amount of heat required to raise the temperature of one gram of water one degree Celsius. The value for water is 4.19 J/(g·°C).

The total energy possessed by a system due to the position and motion of its particles is called *heat.*

$$\text{heat} = E_T = E_P + E_K$$

E_T - total energy
E_P - potential energy
E_K - kinetic energy

The symbol Δ will be used to represent *a change in ....*
If no phase change is involved, no change of potential energy occurs ($\Delta E_P = 0$).
All the change in heat is then due to the change in kinetic energy only.
The heat absorbed or released during a change can be calculated from the mass, change in temperature, and specific heat.

$$\text{heat} = \Delta E_K = ms\Delta T$$

ENERGY

MEASUREMENT OF ENERGY CHANGE

Heat capacity

Manufacturers of calorimeters usually give the *heat capacity*, in joules per degree Celsius for each calorimeter. The heat capacity is the amount of heat required to raise the temperature of the calorimeter and its contents one degree Celsius. The heat capacity is the ms of the formula:

$$heat = ms\Delta t.$$

Example 1:

The heat capacity of a calorimeter is listed as 364 kJ/°C. The calorimeter was used for an experiment and a temperature rise of 10.4°C was observed. Calculate the energy released by the change.

$$\begin{aligned} \text{heat} &= ms\Delta t \\ &= 364\ \text{kJ/°C} \times 10.4\text{°C} \\ &= 3786\ \text{kJ} \\ &= 3.79\times10^{3}\ \text{kJ} \end{aligned}$$

Example 2:

A calorimeter is used as a hot water bath to warm an acid solution. The temperature of the 500 g of water in the calorimeter decreases from 90.0°C to 78.0°C. How much heat was used to heat the acid solution?

$$\begin{aligned} heat &= ms\Delta t \\ &= 500\ \text{g} \times 4.19\ \text{J/(g·°C)} \times 12.0\text{°C} \\ &= 25\ 140\ \text{J} \\ &= 2.51 \times 10^{4}\ \text{J or } 25.1\ \text{kJ} \end{aligned}$$

Example 3:

If the original temperature of the acid in Example 2 was 20.0°C and its specific heat is 4.50 J/(g·°C), calculate the mass of acid heated in the calorimeter.

Assuming no heat loss from the system, the 2.51×10^4 J of energy lost by the water must have been gained by the acid (conservation of energy).

$$\begin{aligned} heat &= ms\Delta t \\ m &= \frac{heat}{s\Delta t} \\ &= \frac{2.51 \times 10^{4}\ \text{J}}{4.50\ \text{J/(g·°C)} \times 58\text{°C}} \\ &= 96.2\ \text{g} \end{aligned}$$

96.2 g of acid were heated in the calorimeter.

Question:

1. How much heat is required to raise the temperature of 250 g of water 12.0°C?

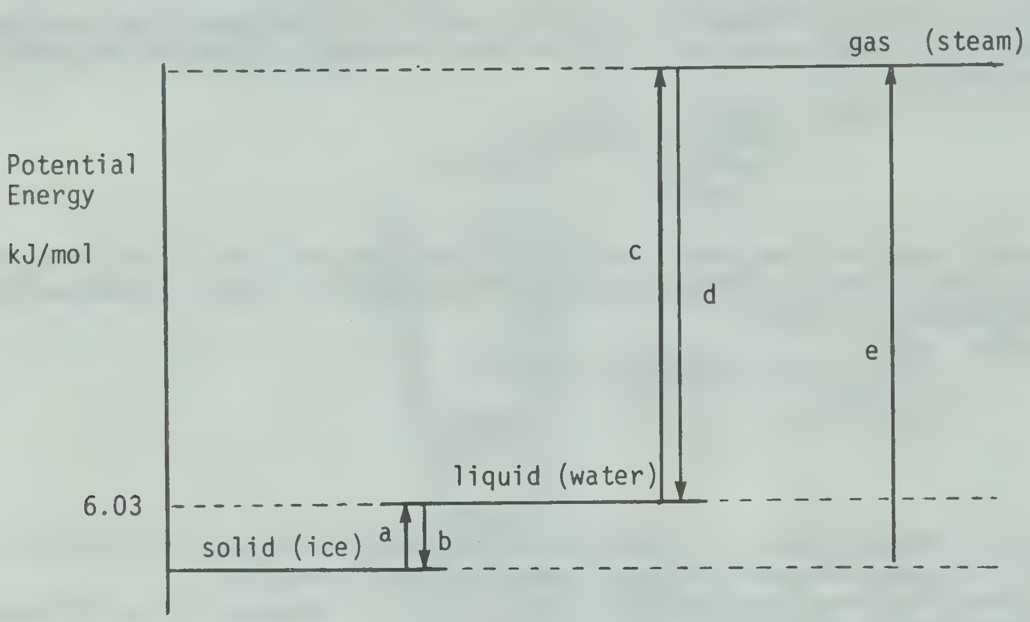
2. A student is required to warm 2.02 kg of a sugar solution from room temperature, 22.2°C, to 75.0°C. The specific heat of the solution is 5.21 J/(g·°C). How much heat would be required?

PHASE CHANGES

Energy Associated with Phase Change

During a phase change such as melting or vaporization, the energy added does work against the *inter*molecular forces and is stored as *potential energy*. The average *kinetic energy* does not change and therefore the temperature does not change when a phase change occurs. During freezing and condensation the process is reversed and the stored energy is released.

Figure L3 shows how potential energy increases as ice is converted to water and finally to steam. Identify the various phase changes by their proper name.



Energy of Phase Changes
Figure L3

Consider the *continuous* heating of a piece of ice at -10°C in a closed system, until the system contains steam at 110°C. As heat is added the ice absorbs energy and the temperature of the ice starts to rise (the kinetic energy of the ice has increased). At 0°C the heat added does work to overcome the intermolecular attractive forces in the ice crystal, and is stored as *potential energy*. During this phase change no change in *temperature* occurs and therefore the average *kinetic energy* does not change, The energy required for the phase change without a temperature change is called the *molar heat of fusion*. For ice the molar heat of fusion is 6.03 kJ/mol. When all the ice has melted the added heat increases the temperature (and hence kinetic energy) of the water molecules to 100°C. At 100°C the added heat does work to overcome the intermolecular attractive forces operating between water molecules, and moves them apart to assume the molecular distribution of a gas. Once again the extra energy is stored as potential energy (see Figure L3). At this point the system contains steam at 100°C. The energy required for the liquid to gas phase change without a change in temperature is called the *molar heat of vaporization*. For water the value is 40.8 kJ/mol. No change in temperature or kinetic energy takes place until all water molecules have been changed to steam molecules at 100°C. Now the steam molecules absorb energy causing the temperature to rise to 110°C and the kinetic energy to increase.

MOLAR HEAT

Introduction

In the section of this unit on *calorimetry* the formula $heat = ms\Delta t$ was developed. The amount of heat produced by a phase change is calculated using this formula (see Example 1 which follows).

Example 1:

A *small amount* of wax is poured into the reaction vessel of a calorimeter containing 500 ml of water and allowed to solidify. How much heat was absorbed by the water if the temperature increased 1.00°C?

$$\begin{aligned} heat &= ms\Delta t \\ &= 500\ \text{g} \times 4.19\ \text{J/(g}\cdot\text{°C)} \times 1.00\text{°C} \\ &= 2100\ \text{J} \\ &= 2.10\ \text{kJ} \end{aligned}$$

210 kJ of heat are released by the wax when it solidifies. The assumption is made that all 2.10 kJ are absorbed by the water in the calorimeter. However, no mention was made in this example of *how much* wax solidified. If the mass of the wax was given the *molar heat of solidification* of wax could be calculated; i.e., the amount of heat released when one mole of wax solidifies.

Calculating Molar Heat of a Phase Change

The *molar heat* of a substance is the amount of heat absorbed or released *per mole* of the substance when it undergoes a change. Molar heat can be calculated using the formula,

$$molar\ heat = \frac{heat}{\#\ moles} \quad \text{in units of kJ/mol.}$$

Example 2:

Suppose, in Example 1, 35.2 g of wax (assume wax is $C_{25}H_{52}$) are used; calculate the *molar heat of solidification* of wax.

(from Example 1 heat = 2.10 kJ)

$$\begin{aligned} \#\ \text{moles}_{\text{wax}} &= \frac{\text{mass}}{\text{molar mass}} \\ &= \frac{35.2\ \text{g}}{352\ \text{g/mol}} \\ &= 0.100\ \text{mol} \end{aligned}$$

$$\begin{aligned} \text{molar heat} &= \frac{\text{heat}}{\#\ \text{moles}} \\ &= \frac{2.10\ \text{kJ}}{0.100\ \text{mol}} \\ &= 21.0\ \text{kJ/mol} \end{aligned}$$

The molar heat of solidification of wax is 21.0 kJ/mol.

Note:

The molar heats of other phase changes such as melting, vaporization, condensation, and sublimation can be calculated in exactly the same manner. In addition to calculating heat and molar heat, the formulas can be used to calculate mass, molar mass, change in temperature, and number of moles.

ENERGY

MOLAR HEAT

1. A 38.4 g piece of solid iodine is dropped into the reaction vessel of a calorimeter. When all the iodine has *sublimed*, the temperature of the 500 g of water decreased by 4.50°C. Calculate the *molar heat of sublimation* for iodine.

2. Several droplets of a molten metal are dropped into 250 ml of water at 24.0°C and an increase in temperature of the water of 12.0°C is observed. The molar mass of the metal is 100 g and its molar heat of solidification is 252 kJ/mol. Calculate the mass of the metal dropped into the water.

3. A calorimeter contains 300 ml of water at 94.0°C. When 4.00 g of an unknown pure solid substance is dropped into the reaction vessel of a calorimeter, the solid melts when the temperature of the water has dropped to 34.0°C. The molar heat of melting of the solid is 3.00×10^3 kJ/mol. Calculate the molar mass of the solid.

MOLAR HEAT

4. In the reaction vessel of a calorimeter 2.00 kg of steam at 100°C condenses to 2.00 kg of water at 100°C. If the *molar heat of condensation* of steam is 40.8 kJ/mol, calculate the change in temperature of the 400 ml of water in the calorimeter.

5. How many moles of liquid sodium chloride solidify in a calorimeter to raise the temperature of 200 ml of water by 3.20°C? The *molar heat of solidification* of sodium chloride is 28.5 kJ/mol.

6. Calculate the total amount of energy required to convert 54.0 g of ice at -20°C to 54.0 g of steam at 150°C.

Note:

The molar heats of chemical changes are calculated in exactly the same way. Questions of this type will be found in the section on Chemical Change.

PHASE CHANGE

Calorimetric Determination of the Molar Heat of a Phase Change - Lab L1

Purpose:

Part A: To determine the molar heat of a phase change.

Part B: (Optional) To determine the *heat capacity* of the calorimeter (styrofoam cups and thermometer).

Materials:

- 1 - 18 x 150 mm test tube partially filled with wax
- 1 - 18 x 150 mm test tube
- 1 - 250 mℓ beaker
- 2 - nested styrofoam cups
- 1 - styrofoam cup for optional part
- 1 - test tube holder
- 1 - thermometer (-10°C to 110°C)
- 1 - 10 cm ring
- 1 - ringstand
- 1 - bunsen burner
- 1 - wire gauze

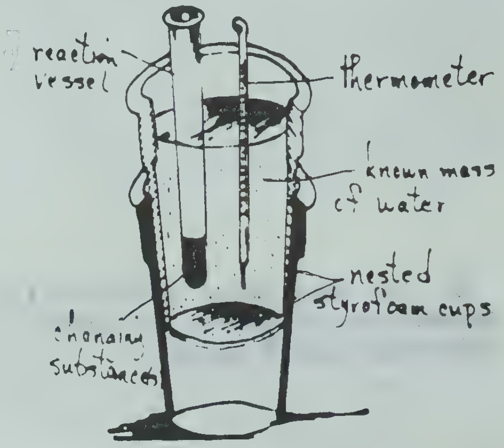


Figure L6

Note:

If directed by the teacher to do the optional part of the experiment first, refer to the optional section at the end of the experiment.

Procedure:

Part A:

1. Fill the 250 mℓ beaker with about 200 mℓ of water and bring the water to a boil.
2. Determine and record the mass of the 18 x 150 mm test tube plus wax.
3. Determine and record the mass of the empty 18 x 150 mm test tube.
4. Determine and record the mass of the two nested styrofoam cups.
5. Fill the two nested styrofoam cups half full of tap water.
6. Adjust the temperature of the water to about 2°C below room temperature using hot and cold water.
7. Determine and record the mass of the nested styrofoam cups and water. Return the thermometer to the water.
8. Using a test tube holder, place the test tube with wax into the boiling water until the wax is *just* melted.
9. Remove the test tube with wax from the hot water and cool in air until a faint cloudiness appears in the wax. Do Step 10 while the cooling occurs.
10. Determine and record the temperature of the water in the nested styrofoam cups.
11. Place the test tube plus wax into the water in the nested styrofoam cups as soon as cloudiness appears.
12. Stir the cold water bath and record the highest temperature reached by the water bath.
13. Repeat the experiment if the time permits, using a slightly *larger* mass of wax.

PHASE CHANGE

Observations:

	Trial 1	Trial 2
mass of test tube plus wax		
mass of empty test tube		
mass of styrofoam cups plus cold water		
mass of styrofoam cups		
initial temperature of water		
final temperature of water		

Preliminary Calculations:

	Trial 1	Trial 2
mass of wax		
mass of cold water		
temperature change		

Calculations:

If Part B, the optional part of the lab was done, do Part B Calculations first.*

Determine the molar heat of solidification of candle wax. Assume the candle wax is $C_{25}H_{52}$.

Trial 1 Trial 2

* total heat = msΔt + heat capacity x Δt
= heat gained by water + hear gained by calorimeter.

PHASE CHANGE

Part B: Optional determination of the *heat capacity* of the calorimeter.

Procedure:

Part B:

1. Add 90.0 g of H_2O, at approximately room temperature, to two nested styrofoam cups, which will serve as the calorimeter.
2. Record the temperature of the water to the nearest 0.1°C.
3. Obtain 90.0 g of H_2O, at approximately 10°C higher than room temperature, in another styrofoam cup.
4. Record the temperature of the warm water to the nearest 0.1°C.
5. Add the warm water to the cool water in the two nested styrofoam cups and stir continuously.
6. Record the maximum temperature.

Calculations:

1. Calculate the heat gained by the cool water.

2. Calculate the heat lost by the warm water.

3. Calculate the difference between the heat gained and the heat lost. The difference is the heat absorbed by the styrofoam cup, the thermometer, and the surroundings.

4. Calculate the *heat capacity* (mass x specific heat, *ms*) of the calorimeter.

$$\textit{heat capacity} = \frac{\text{heat absorbed by the cup and thermometer}}{\text{rise in temperature}}$$

$$\textit{heat capacity} = \frac{heat}{\Delta t} = ms$$

PHASE CHANGE

Questions:

1. What assumptions were made in setting up the lab and in doing the calculations?

2. Why were nested styrofoam cups used instead of a beaker to hold the cold water?

3. Would the expected results obtained in this lab be higher or lower than the accepted value? Explain.

4. Why is the temperature of the water adjusted to room temperature minus 2.0°C?

5. Devise an experiment which would eliminate some of the sources of error in this lab.

CHEMICAL CHANGE

Introduction

A chemical change is a transformation involving an energy change in which one substance is converted into another substance.

At the present time the energy released during a chemical reaction is man's prime method of converting available energy to useful energy. Examples of such uses of exothermic chemical reactions are:

1. steam generating stations where fossil fuels are burned to produce electricity (e.g., coal at Lake Wabamum, Alberta or natural gas at Edmonton, Alberta).
2. motor vehicles where gasoline or diesel fuel is burned to produce mechanical energy
3. heating of homes where fossil fuels such as natural gas, propane, oil, coal and wood are burned to produce heat energy
4. batteries, where chemicals react to produce electricity directly
5. living organisms, including man, where chemical energy, stored in food, is used to maintain body temperature and to promote growth.

Energy Stored in Bonds

Energy added to matter during a *phase change* is stored as potential energy in the particles that make up that sample of matter. Work is done against the *intermolecular forces* (bonds) that hold matter together. The phase change can be represented by equations such as

$$\text{matter (solid)} + \text{energy} \rightleftharpoons \text{matter (liquid)}$$

The rearrangement of chemical bonds also involves energy. Energy must be added (work done) to separate the two particles held together by a *chemical bond*. This energy is stored as potential energy in the separated particles and is released when the particles are allowed to come together to form a new bond.

The energy required to break a bond, which is also the energy released when a bond is formed, is called *bond energy*. Consider the reaction:

$$H_{2(g)} + Cl_{2(g)} \longrightarrow 2HCl_{(g)} + 184\ \text{kJ}$$

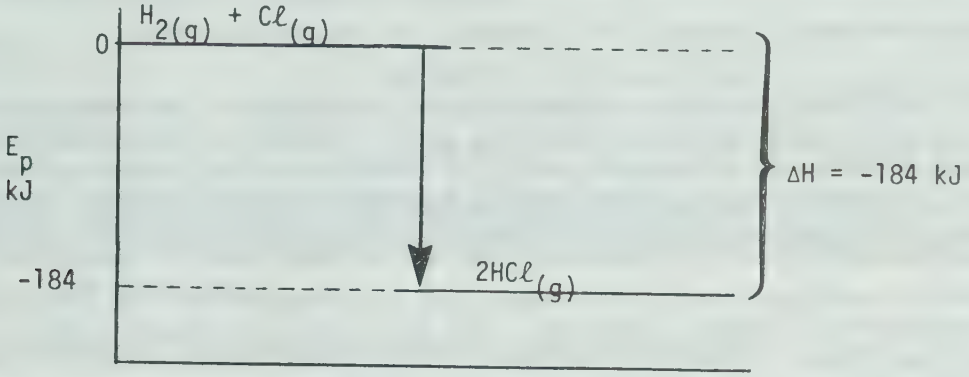
Energy is required to break one mole of H-H and one mole of Cl-Cl bond and is stored as potential energy in the separate particles. Energy is released when two moles of H-Cl bonds form. Quantitatively it can be seen that 184 kJ more energy was released when the new bonds were formed than was absorbed when the old bonds were broken. The bond energies and net energy of the reaction are summarized in Table L2.

Table L2
Bond Energies in a Chemical Change

Energy Absorbed by the Reaction	Energy Released by the Reaction
1 mol H-H bond energies	2 mol H-Cl bond energies
+ 1 mol Cℓ-Cℓ bond energies	______________
total energy absorbed	total energy released
net energy change = total energy released - total energy absorbed	

CHEMICAL CHANGE

The energy change that is observed in any reaction is the difference between the energy released when new bonds are formed and the energy absorbed when the original bonds are broken. The reaction of hydrogen and chlorine gases is represented graphically in Figure L



Energy Change of Combustion of $H_{2(g)}$ with $Cl_{2(g)}$
Figure L7

Notice in Figure L7 that the reactants start with a certain amount of energy and the product ends up with 184 kJ *less* energy. There has been a net energy loss from the system of 184 kJ. The ΔH *notation* and the negative value for ΔH are discussed next.

Questions:

1. Define *bond energy*.

2. The *Sherritt-Gordon Refinery* at Fort Saskatchewan uses the *Haber Process* to synthesize ammonia, NH_3. Using the equation for the Haber Process written below, describe the relationship between bond energies and net energy of reaction (heat of reaction, ΔH).

$$N_{2(g)} + 3H_{2(g)} \longrightarrow 2NH_{3(g)} + 101 \text{ kJ}$$

CHEMICAL CHANGE

Representing Energy Change in a Chemical Reaction

The energy change in a chemical reaction can be represented in one of two ways.

1. Include the energy change as an integral part of the equation.

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(g)} + 484\ \text{kJ}$$

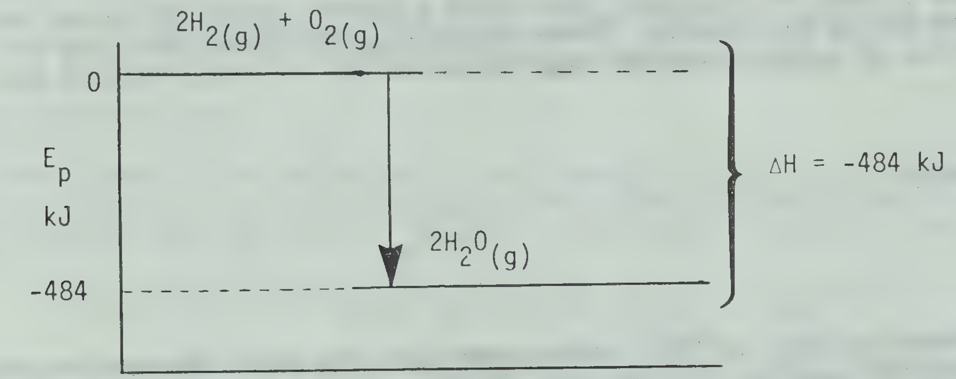
$$N_{2(g)} + 2O_{2(g)} + 33.9\ \text{kJ} \longrightarrow 2NO_{2(g)}$$

The first equation represents an exothermic reaction evolving 484 kJ of heat. The second equation represents an endothermic reaction requiring 33.9 kJ of energy.

2. Use the *ΔH notation.* ΔH is the symbol used to represent the *net* amount of heat released or absorbed by a chemical reaction. The value of ΔH is negative for exothermic reactions (see Figure L7 and L8), and positive for endothermic reactions (see Figure L9).

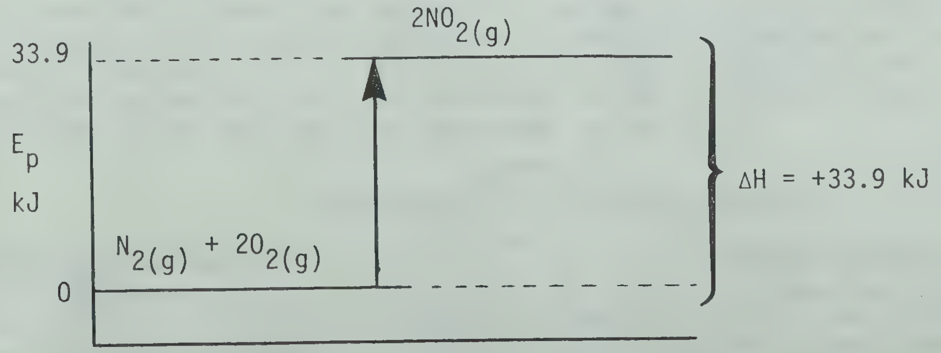
$$2H_2(g) + O_2(g) \longrightarrow 2H_2O(g) \qquad \Delta H = -\ 484\ \text{kJ}$$

$$N_2(g) + 2O_2(g) \longrightarrow 2NO_2(g) \qquad \Delta H = +\ 33.9\ \text{kJ}$$



Exothermic Combustion of $H_{2\ (g)}$
Figure L8

As the combustion of $H_{2(g)}$ proceeds the potential energy, E_p, of the chemicals *decreases*. To represent the decrease in potential energy in Figure L8 the ΔH value is written as a *negative* number.



Endothermic Combustion of $N_{2(g)}$
Figure L9

As the combustion of $N_{2(g)}$ proceeds the potential energy of the chemicals *increases*. To represent the increase in potential energy in Figure L9 the ΔH value is written as a *positive* number.

MOLAR HEAT OF A CHEMICAL CHANGE - LAB L2

Purpose:

To determine the molar heat of a chemical reaction.

Prelab Information:

The heat of reaction will be measured for the neutralization of 1.00 M HCl with 1.00 M NaOH using two nested styrofoam cups as a simple but efficient calorimeter. The heat produced by the neutralization reaction will heat a measured amount of aqueous solution in the calorimeter. Unless directed otherwise by the teacher, the assumption is made that small losses to the calorimeter, air and the thermometer will be ignored. (If directed by the teacher to calculate the heat capacity of the calorimeter refer to the optional part of Lab L1 Page L for details.)

Two further assumption's are made to simplify the calculations.

1. The 1.00 M solutions are considered to be water, with a specific heat of 4.19 J/(g.°C).

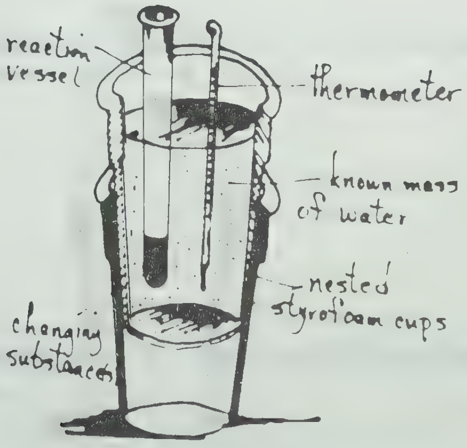
2. One millilitre of water has a mass of very nearly one gram. Therefore the *mass* of the solutions heated, *in grams*, can be assumed to be numerically equal to the *volume, in millilitres*.

Prelab Exercises:

1. Write a balanced chemical equation for the reaction of HCl and NaOH solutions.

2. Why are *equal* volumes and *equal* concentrations of the NaOH and HCl solutions used in this experiment?

3. How many moles of each reactant are used?



A Simple Calorimeter
Figure L10

Materials:

2 - styrofoam cups
1 - 50 ml graduated cylinder
1 - Celsius thermometer
50 ml - 1.00 M HCℓ
50 ml - 1.00 M NaOH
1 - wash bottle (distilled water)

MOLAR HEAT OF A CHEMICAL CHANGE - LAB L2

Procedure:

1. Nest the two styrofoam cups.
2. Measure 50.0 ml of 1.00 M HCl with a graduated cylinder and pour it into the nested styrofoam cups.
3. Using a rinsed 50 ml graduated cylinder, measure 50.0 ml of 1.00 M NaOH and leave it in the graduated cylinder.
4. Record the temperature of both the HCl and NaOH solutions to the nearest 0.1°C and proceed quickly to the next step.
5. Pour the NaOH solution into the HCl solution in the styrofoam cup and carefully stir the mixture with the thermometer.
6. Record the highest temperature obtained, to the nearest 0.1°C

Observations:

initial temperature of HCl solution ________________

initial temperature of NaOH solution ______________

initial average temperature of solutions _____________

final temperature of solution ____________

Calculations:

Calculate the molar heat of reaction. (Refer to Page L 13)

MOLAR HEAT OF A CHEMICAL CHANGE - LAB L2

Conclusion:

$NaOH_{(aq)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(\ell)}$ ΔH = ____________

Questions:

1. What assumptions were made in setting up this experiment and doing the calculations?

2. Explain whether the heat of reaction obtained in this experiment should be higher or lower than the accepted value.

3. What is the purpose of nesting two styrofoam cups to serve as the calorimeter?

4. If a 250 ml pyrex beaker were used as the calorimeter in place of the styrofoam cups would the value of ΔH be larger or smaller? Explain.

5. Calculate the percent error. The heat of reaction for a strong acid-strong base neutralization is -56.6 kJ/mol of water formed.

CHEMICAL CHANGE

Applications of Heat of Reaction

1. A propane hot water heater is used to heat 200 ℓ of water from 25.0°C to 78.2°C. Calculate the amount of heat consumed.

2. If 1.00 g of propane were used to heat 200 ℓ of water from 25.0°C to 78.2°C (see Question 1), calculate the molar heat of combustion of propane.

3. The instruction manual for a small propane camp stove indicates that the stove will yield 1000 kJ of usable energy per mole of propane consumed under average conditions. Camper wishes to heat 1.00 ℓ of water from 27.0°C to 95.0°C in order to make soup. What mass of propane was consumed?

4. In a calorimetry experiment the burning of 5.08 g of benzene released enough heat to raise the temperature of 5.00 kg of water from 10.1°C to 19.6°C. Calculate the molar heat of combustion of benzene.

ENERGY

CHEMICAL CHANGE

Heat of Formation, H

Heat of formation is the change in potential energy which occurs when a compound is formed from the elements. All elements in their natural state are assumed to have a heat of formation (H) of zero. The heats of formation for some common substances are included on the *ALCHEM* data sheet.

Heats of formation are useful for predicting the heat of a reaction. It has been verified experimentally that the heat of a reaction can be obtained by subtracting the sum of the heats of formation of the reactants from the sum of the heats of formation of the products; i.e.,

$\Delta H = \Sigma H_P - \Sigma H_R$, where ΣH_P and ΣH_R represent the sums of the heats of formation of products and reactants respectively. The general method of obtaining the heat of reaction from the heats of formation is summarized below.

Step 1: Write the balanced equation for the reaction.

Step 2: Obtain the heats of formation of the reactants and products from the ALCHEM data sheet, and write them down under the equation.

Step 3: Multiply the heat of formation values by the coefficients from the balanced chemical equation.

Step 4: Sum the heats of formation of the reactants and products separately.

Step 5: Substitute the sums into the formula; $\Delta H = \Sigma H_P - \Sigma H_R$.

Step 6: Write the heat of reaction in ΔH notation with units of kJ *for the reaction as written.*

Example:

Predict the heat of reaction for the combustion of one mole of ethane; i.e., predict the molar hear of combustion of ethane. *For molar heats of reaction always balance the equation using a coefficient of one for the important reactant or product.*

Step 1:	$C_2H_{6(g)}$	$+ \frac{7}{2} O_{2(g)}$	$\longrightarrow$	$2CO_{2(g)}$	$+ 3H_2O_{(g)}$
Step 2:	-84.6	0		-393.6	-242
Step 3:	1(-84.6)	$\frac{7}{2}(0)$		2(-393.6)	3(-242)
Step 4:	-84.6			-787.2	-726.0
				-1513.2	

Step 5: $\Delta H = \Sigma H_P - \Sigma H_R$

$= (-1513.2) - (-84.6)$

$= -1428.6$

Step 6: $\Delta H = -1428.6$ kJ

The heat of reaction for the combustion of one mole of ethane (the molar heat of combustion of ethane) is 1428.6 kJ.

CHEMICAL CHANGE

Using Heats of Formation (H) to Calculate Heats of Reaction (ΔH)

1. The Sheritt-Gordon Nickel Refinery at Fort Saskatchewan synthesizes hundreds of tonnes of ammonia, NH_3, in a reaction called the *Haber Process*. Nitrogen gas and hydrogen gas are reacted at high temperature and pressure to produce ammonia. Write a balanced chemical equation for the reaction and calculate the molar heat of reaction for ammonia from the heats of formation.

2. Water is electrolyzed by passing a direct current of electricity through water to which a small amount of sulfuric acid has been added. (The purpose of the acid will be studied in ALCHEM Unit M, *Electrochemistry*.) Complete the balancing of the equation and calculate the molar heat of reaction of water.

$$H_2O_{(\ell)} \longrightarrow H_{2(g)} + O_{2(g)}$$

3. The very rapid, almost explosive, reaction of aluminum and iron(III) oxide is called the *Thermit reaction*. Due to the speed of the reaction and the very large amount of energy released when the iron is produced, the Thermit reaction was once used for on-site welding.

$$Al_{(s)} + Fe_2O_{3(s)} \longrightarrow Al_2O_{3(s)} + Fe_{(s)}$$

Balance the equation and calculate the heat of reaction for two moles of aluminum.

CHEMICAL CHANGE

4. Carbohydrates, such as sugars and starches, are oxidized in the body to provide needed energy. The oxidation of carbohydrates is the reverse of photosynthesis (see Figure L2) and therefore the energy released in respiration must ultimately have come from the sun (solar energy). The heat of formation for glucose is -900 kJ/mol. Calculate the heat of combustion for glucose and compare with the ΔH for the photosynthesis reaction written above Figure L2.

5. Sucrose is table sugar, $C_{12}H_{22}O_{11}$, a major source of energy in the food consumed by man. The heat of formation for sucrose is -1742 kJ/mol. Calculate the molar heat of combustion of sucrose, Assume the products of combustion are carbon dioxide and water vapor.

6. A housewife, in reaching across a boiling kettle for a dish, receives second degree burns to her arm from the steam. Assuming 0.500 g of steam condensed upon contact with her arm, how much energy caused the burn?

CHEMICAL CHANGE

Solving Problems Involving Heats of Reaction

Example:
How much methanol would be burned to heat 300 mℓ of water from room temperature (23.0°C) to the boiling point (98.0°C in Alberta).

Step 1: Determine the heat of reaction as shown earlier.

$$CH_3OH_{(\ell)} + \frac{3}{2}O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)}$$

$$(-238.6) + \frac{3}{2}(0) \neq (-393.6) + 2(-242.0)$$

$$-238.6 \neq -877.6$$

$$\begin{aligned} \Delta H &= \Sigma H_P - \Sigma H_R \\ &= (-877.6) - (-238.6) \\ &= -639.0 \text{ kJ} \end{aligned}$$

Step 2: Determine the amount of heat required to be absorbed by the water.

$$\begin{aligned} heat &= ms\Delta t \\ &= 300 \text{ g} \times 4.19 \text{ J/(g}\cdot\text{°C)} \times 75.0\text{°C} \\ &= 94\ 203 \text{ J} \\ &= 94.2 \text{ kJ} \end{aligned}$$

Step 3: Solve for the number of moles of methane ($\#moles = \frac{heat}{molar\ heat}$).

$$\begin{aligned} \#moles_{CH_3OH} &= \frac{heat}{molar\ heat} \\ &= \frac{94.2 \text{ kJ}}{639.0 \text{ kJ/mol}} \\ &= 0.0147 \text{ mol} \end{aligned}$$

Step 4: Determine the mass of methanol required.

$$\begin{aligned} mass_{CH_3OH} &= \#moles \times molar\ mass \\ &= 0.0147 \text{ mol} \times 32.0 \text{ g/mol} \\ &= 0.470 \text{ g} \end{aligned}$$

The mass of methanol required is 0.470 g.

CHEMICAL CHANGE

Heats of Reaction

1. Methane, one of the main constituents of natural gas, is very difficult to liquify. Therefore, it is very expensive to transport by boat. Methane can be oxidized to methanol fairly inexpensively. Methanol, a liquid, is cheap to transport. Compare the molar heats of combustion for methane and methanol and thus their value as fuels.

2. Methanol is a commonly used fondue fuel. How much methanol would be required to heat 1.25 ℓ of a processed hydrogenated soybean oil from 22.0°C to 97.0°C. (Hydrogenation is discussed in *Unit O, Foods and Their Analogs.*) Assume the oil has a specific heat of 8.90 J/(g·°C) and a density of 0.880 g/ml.

CHEMICAL CHANGE

1. How much more energy would a person derive from drinking a cup of coffee containing [illegible] g of dissolved sucrose than he would derive from coffee without any sugar? Assume all 5.00 g of sucrose consumed are oxidized in the body cells to produce energy.

2. [illegible] designed *Bic lighter* contains 9.80 g of liquid butane. An average use of the lighter consumes 59.0 J of energy. The products of the combustion are carbon dioxide and water vapor. Calculate the number of times the owner may *flick his BIC* before [illegible] has been spent.

CHEMICAL CHANGE

Questions 5 and 6 refer to the following information.
In an average year a 120 m^2 bungalow will use 3.00×10^8 kJ of heat.
Assume natural gas is pure methane, and burns to produce water vapor

5. Determine the number of moles of methane required to heat this bungalow for one year.

6. If 1.00 moles of natural gas costs 6.00×10^{-4} dollars, determine the cost of heating the house for one year.

7. A 5 m x 9 m enclosed swimming pool requires an average of 2.10×10^5 kJ per hour to heat the pool in cold weather. Determine the cost of operating this pool for one month (30 days) in winter. Natural gas costs \$0.060/mol.

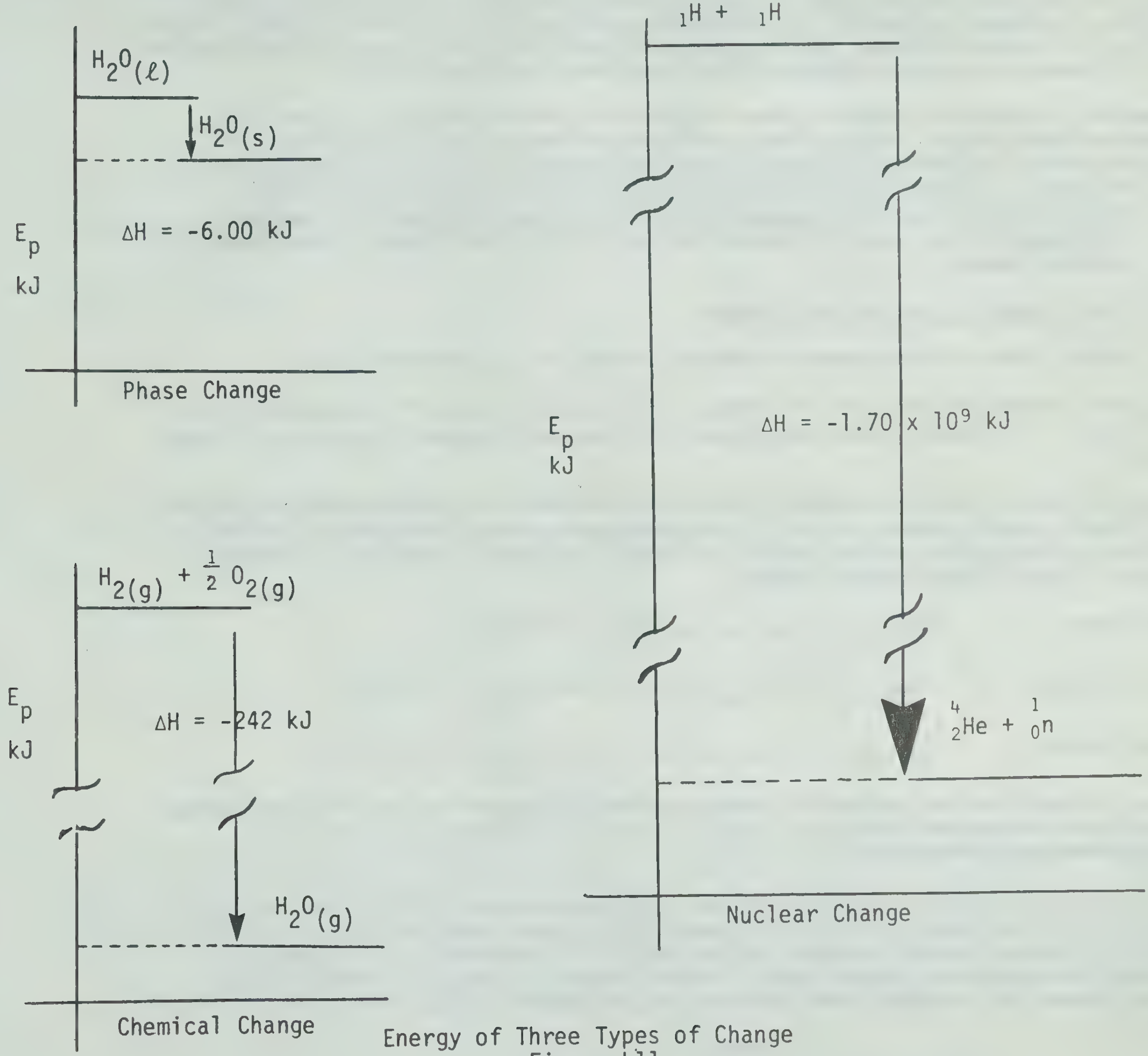
CHEMICAL CHANGE

8. A hot water heater requires about 3500 MJ to operate for one month. What would the natural gas cost? Natural gas costs $6.00 x 10^{-4}/mol.

9. Nitromethane, CH_3NO_2, is an exotic fuel used by the fuel class dragsters for drag racing. One of the top dragsters in the world is owned by Ray Peets and driven by Gary Beck, (1974 World Champion) both Edmontonians. Although nitromethane has a lower heat of reaction than gasoline, which can be assumed to be an isomer of octane called 2,2,4-trimethylpentane, the fact that it produces the energy much faster makes it a very suitable fuel for dragging. Compare the molar heats of combustion of nitromethane and gasoline. Assume combustion of nitromethane yields carbon dioxide, nitrogen dioxide, and water vapor.

COMPARING PHASE, CHEMICAL, AND NUCLEAR CHANGES

Figure L11 graphically represents the different amounts of energy involved in the three types of changes



Energy of Three Types of Change
Figure L11

Obviously the three graphs are not drawn to the same scale. Assume the distance between the $H_2O_{(\ell)}$ and $H_2O_{(s)}$ on the graph is 1.00 cm (roughly the width of your little finger nail). To draw the other graphs to scale the distance between reactants and products for the chemical change would need to be 40 cm (about twice the width of this page). The nuclear change would require a graph extending from Edmonton, Alberta to a distance of 283 million cm (2830 km). Hopefully this analogy will place the fantastic amount of energy obtainable from a nuclear reaction in its true perspective.

The relative amounts of energy involved in phase, chemical and nuclear changes are summarized in Table L3.

NUCLEAR CHANGE

Introduction

Man has relied almost exclusively on chemical energy throughout most of history. With the population explosion and the ever increasing demand by modern society for still more mechanization the demand for energy is threatening to exhaust known sources of fossil fuels-natural gas, coal, and petroleum. Although new sources of fossil fuels are being discovered the world's resources are being depleted faster than they are being found.

Fortunately, in the 20th century a new, very powerful form of energy, *nuclear energy* has become available to man. Nuclear energy is derived from nuclear reactions which involve the breaking and forming of bonds between the protons and neutrons within the nuclei of the atoms. Nuclear forces are much stronger than the electrostatic forces operating between positive nuclei and negative electrons (from which chemical energy is derived). Consequently the amount of energy involved in a nuclear change is much greater than in a chemical change.

Comparing Phase, Chemical, and Nuclear Changes

A nuclear change differs from a phase or chemical change in two important ways.

1. In nuclear changes new elements are formed. One element is actually transformed into another element. For example radium may change into lead and hydrogen isotopes may change into helium.

2. In general nuclear changes produce millions of times more energy than chemical changes. Chemical changes *usually* produce 10 to 10^2 times more energy than phase changes, although there is some *overlap* as the circled values in Table L3 indicate. Consider and compare the energies of the following three changes.

1. phase change $H_2O_{(\ell)} \longrightarrow H_2O_{(s)} + 6.00 \text{ kJ}$

2. chemical change $H_{2(g)} + 1/2\ O_{(g)} \longrightarrow H_2O_{(g)} + 242 \text{ kJ}$

* 3 nuclear change ${}^2_1H + {}^3_1H \longrightarrow {}^4_2He + {}^1_0n + 1.70 \times 10^9 \text{ kJ}$

The nuclear change shows the isotope of hydrogen called *deuterium* reacting with the isotope called *tritium* to produce helium and a neutron. One billion, seven hundred million kilojoules of energy are released.

* *Note: The writing of nuclear reaction equations will be studied in ALCHEM Elective Unit LL, Nuclear Energy and Radiation, which follows directly after ALCHEM Unit L. In addition ALCHEM Unit LL studies Canada's nuclear reactor CANDU. The important role played by radioisotopes in medicine is another area of major concern. In general, radio isotopes are used as tracers for analysis and diagnosis, or as radiant energy emitters for treatment (therapy).*

Table L3

Comparing Energy of Phase, Chemical, and Nuclear Changes

Change	Equation	Heat of Reaction ΔH	General Range of ΔH
Phase	$H_2O_{(\ell)} \longrightarrow H_2O_{(s)}$	-6.00 kJ	10^1 to $\boxed{10^2}$ kJ
Chemical	$H_{2(g)} + \frac{1}{2}O_{2(g)} \longrightarrow H_2O_{(g)}$	-242 kJ	$\boxed{10^2}$ to 10^3 kJ
Nuclear	${}^2_1H + {}^3_1H \longrightarrow {}^4_2He + {}^1_0n$	-1.70×10^9 kJ	10^6 to 10^9 kJ

ENERGY

OVERVIEW

Phase changes, chemical changes and nuclear changes all involve changes in energy. The following reactions can be used as commercial sources of energy.

1. Calculate the heat released per kilogram of steam condensed.

$$H_2O_{(g)} \longrightarrow H_2O_{(\ell)} \qquad \Delta H = -40.8\ \text{kJ}$$

2. Calculate the heat released per kilogram of hydrogen burned.

$$H_{2(g)} + O_{2(g)} \longrightarrow H_2O_{(\ell)} \qquad \Delta H = -286\ \text{kJ}$$

3. Calculate the heat released per kilogram of U-235 (molar mass of 235 g/mol) which undergoes fission.

$${}^{235}_{92}U + {}^{1}_{0}n \longrightarrow {}^{92}_{36}Kr + {}^{141}_{56}Ba + 3\,{}^{1}_{0}n \qquad \Delta H = -\ 1.90 \times 10^{10}\ \text{kJ}$$

4. How many grams of ice would need to melt in a portable cooler to lower the temperature of six 341 ml bottles of *liquid refreshment* from 31.8°C to 8.0°C. Assume all of the heat gained by the ice was lost by the contents of the bottles. Also assume the contents of the bottle have the same specific heat as water.

$$H_2O_{(s)} \longrightarrow H_2O_{(\ell)} \qquad \Delta H = 6.00\ \text{kJ}$$

OVERVIEW

5. Coal and gasoline are two fuels which are widely used today. Because of dwindling reserves of coal and gasoline,and pollution problems associated with their use, methanol and hydrogen gas have been proposed as fuels of the future. Use the heats of formation given in the *ALCHEM* data sheet to calculate the maximum heat produced per kilogram burned for each of these four fuels. Consider coal to be pure carbon and gasoline to be pure octane and assume any water produced is in the gas phase. What problems might be associated with the use of methanol and hydrogen gas as fuels?

ALTERNATE FORMS OF ENERGY

Introduction

Early in this unit, brief mention was made of a few alternate forms of energy such as electrical, solar, and nuclear energy. This section will further discuss these alternatives and add a few more. In addition, nearly every form of energy production causes undesirable effects upon the enviroment. These effects will be discussed in this section.

The population explosion and the ever increasing demand for more mechanization and for a higher standard of living means that both industrialized and developing nations must have a constantly increasing supply of energy. The problem facing scientists today is how to provide more and more energy every year, and, at the same time, reduce the damage to the environment which is caused by energy production.

Electrical Energy

Electricity is a clean, easy to use source of energy but the generation of electricity causes severe environmental problems in certain localities.

About three times as much energy is expended to produce electricity as is obtained from steam electric power plants. Most of the energy lost is wasted in the form of heat. River or lake water is generally used by steam electric plants for cooling. The hot water is then returned to the river or lake where it causes *thermal pollution*. When lake water is warmed by several degrees the solubility of the dissolved gases decreases. In particular, the amount of dissolved oxygen, which is essential for fish to breath, decreases as the water warms up. As a result, the fish that are consumed by man often die or leave the area in search of a more suitable environment. The hardier, less desirable *trash fish* are left behind. For example, trout thrive in water at 8°C but die when the temperature rises above 26°C. Less desirable fish such as carp can tolerate temperatures of 32°C and higher.

Practically all chemical and biochemical reactions speed up at higher temperatures. Thermal pollution causes the temperature of water to increase and therefore causes other far reaching environmental problems.

1. Fish spawn earlier and their eggs hatch unusually early in the season, often before there is enough to eat.

2. Algae and aquatic plants grow much more quickly, upsetting the *ecological balance* of that particular body of water.

3. Normal body processes increase, including the nervous system. The fish become more active, their heartbeat increases, and they die sooner.

Unfortunately, there are no clean sources of energy widely used today. The *fuel cell*, discussed later in this section is as close to *clean energy* as is currently available.

Hydroelectric Energy

Hydroelectric plants consist simply of a dam across a stream, lake or river. Water falling over the dam turns a series of giant turbines which turn the rotors of electric generators. What kind of environmental problems could this seemingly *clean* method of energy production cause?

First, an enormous mass of water suddenly builds up behind the dam. The sudden increase in downward pressure could give rise to earthquakes in the region, particularly if the region is geologically unstable. Very large earthquakes have been caused in this way.

Second, a dam can have a disastrous effect on the natural environment and the ecological balance of a region. The Aswan High Dam in Egypt was built for flood control as well as for the production of electricity. After the dam was built it was found that the floods had brought with them nutrient containing sediments from the lands upstream. With no more flooding, the flood plains became less fertile. Now chemical fertilizers (which require energy to produce) are needed to restore the fertility to the flood plains.

ALTERNATE FORMS OF ENERGY

A third factor to consider is the large body of *still water* behind the dam which serves as an ideal medium for the growth of mosquitoes, and other disease causing organisms that affect man, sometimes fatally.

Finally, a dam may block the flow of nutrients needed downstream by fish, thus curtailing the fishing industry, sometimes drastically.

These are but a few of the many problems caused by this relatively clean source of energy, hydroelectric energy.

Geothermal Energy

Energy derived from heat contained in rock formations under the surface of the earth is called *geothermal energy*. In areas where active volcanoes are found there are pockets of molten rock found deep in the earth. Surface water seeps down into these pockets and is transformed into steam at very high temperature and pressure. Due to the high pressure the steam forces its way through cracks in the rock formation back to the surface of the earth where it often explodes out of the surface as a *fumarole*. Many of you have probably seen a good example of this in Yellowstone National Park in the state of Wyoming, U.S.A. *Old Faithful* the famous fumarole (not geyser) in Yellowstone erupts approximately every 65 min blowing approximately 50 000 ℓ of water and steam in a magnificant spectacle to a height of 53 m (the height of a 17 storey building).

Man can make use of the geothermal steam energy by drilling holes and running pipes into the pockets containing the steam. The steam can then be used directly for heating or indirectly to generate electricity.

Iceland, Italy, New Zealand, and United States have considerable areas of volcanic activity and are using geothermal energy to a limited extent.

The Geysers Power Plant in northern California is the only power plant in United States using a significant amount of geothermal energy. In 1975 the Geysers plant produced about as much energy as a conventional modern electric plant (2.16 x 10^6 MJ).

Geothermal power plants are not without environmental problems. Along with the steam, various harmful sulfur containing gases may be released. In addition the release of large quantities of steam may have a *settling* effect on the earth's crust. To guard against the possibility of settling, salt water may be pumped into the pockets to replace what gas been taken out.

Plowshare - Geothermal Energy

An interesting project related to geothermal energy is the *Plowshare-Geothermal* project. This project proposes that nuclear explosives be detonated in geothermal hot spots about 1600 m below the surface of the earth. Water is piped down to the fractured layers of hot rock and is converted to steam which is used to operate generators. The steam is then recycled to prevent thermal pollution.

The only significant problem with this form of energy production is the danger of radiation leakage from the nuclear explosions. However, recent experiences with underground nuclear testing indicate that the dangerous radiations can be completely controlled.

Solar Energy

All of our energy comes originally from the sun, except for a small amount of nuclear energy generated here on earth.

Reference back to Figure L2 will indicate that even the energy of fossil fuels can be traced back to the sun. Millions of years ago green plants absorbed solar enery, a chemical reaction (photosynthesis) occurred and carbohydrates were produced (green plants are called *producers*). Animals, including man, are the *consumers*. The consumers ate the plants and thus indirectly took in the solar energy. The plants and animals ultimately died and the decayed remains were buried under millions of tonnes of sediment. The pressure of the sediment slowly, over millions of years, converted the plant and animal remains to fossil fuels.

Sunlight, striking the earth and warming its surface, gives rise to hurricanes,

ALTENATE FORMS OF ENERGY

cloudbursts, blizzards and thunderstorms (the energy of wind, rain, and snow).

Very little use is made of the enormous amount of solar energy that falls on Earth's surface. In some cloudless areas of the world, nearly 3.60 X 10^6 joules of energy per square metre per day strike the surface during daylight hours. So far the major drawback to utilizing solar energy has been the development of inexpensive but effective collectors to absorb the solar energy. However comprehensive research programs may soon be set up to develop effective collectors.

Solar Energy Heats a Swimming Pool

An Edmonton, Alberta resident heats his backyard swimming pool with solar energy. The collectors are equivalent to about three quarters the surface area of the pool and are mounted on top of a car-port at a slope of about 10 degrees. Ideally the collectors should have a surface area equal to the surface area of the pool. The slope of 10 degrees is ideal for the summer months in Alberta.

Each collector panel is about 1 m x 3 m in size with a flat iron bottom and a corrigated sheet iron top. The unit is sealed, with water entering at the bottom of the slope and leaving at the top. The water is pumped through the collectors with a one-twelfth horsepower aquarium motor at a rate of 500 litres per hour.

The maximum temperature the pool water has reached has been 30°C, with the water coming out of the collector at 39°C.

The Research Council of Alberta has been watching the operation of the pool and plans to install special instruments to gather more data. The Research Council has been investigating the use of solar energy and wind energy and in conjuction with other Alberta Government departments will be publishing information on the potential use of solar energy and wind energy in Alberta in late 1976 or early 1977.

For more information and a picture of the pool and the solar energy collector, refer to page 33 of the July 12, 1976 edition of Edmonton's daily newspaper.

Solar Batteries

Another method of harnessing the energy of the sun is to use a device known as a *solar battery* which converts sunlight directly into electricity. The battery consists of a semiconductor such as silicon upon which sunlight is allowed to fall. As the solar energy strikes the silicon crystals, the electrons start to move and electricity is generated.

Solar batteries, used in unmanned space satellites, have continued to power cameras and communication equipment for years without deterioration. Unfortunately, solar batteries are too expensive to be used on earth except in some unmanned weather stations in remote areas.

Someday solar batteries will provide a practical source of clean electrical power.

Nuclear Energy

Nuclear fusion and *nuclear fission* reactions will be studied in detail in an elective unit, ALCHEM Unit LL. The enormous amount of energy achievable, and the related environmental problems associated with nuclear reactions will be discussed in Unit LL.

Fuel Cells

A fuel cell is a possible source of clean energy. It is a device that transforms chemical energy into electrical energy directly. Fuel cells provided the electricity for the American Apollo spacecrafts.

The electrolysis of water requires energy to produce hydrogen and oxygen gases. Electrolysis is discussed in detail in ALCHEM Unit M, *Electrochemistry*. The reverse of *this reaction*, where hydrogen gas combines with oxygen gas, releases the same amount of energy. A cell of this type is called a *fuel cell*. The hydrogen and oxygen are combined at low temperature without combustion and the energy released is electrical energy.

$$2H_{2(g)} + O_{2(g)} \longrightarrow 2H_2O_{(\ell)} + \text{electrical energy}$$

Burning the two gases and using the energy to make steam to generate electricity would be less efficient and would result in thermal pollution. Fuel cells are used to provide electrical energy for a variety of spacecraft applications, and as a by-product of the reaction, drinking water is produced.

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